Worksheet: Metric

Conversion: To do these, you need to know the Metric-Metric conversion scale.

1) 3.786 m = _____km

2) .0083 hg = _____dg

3) 1001 ul = _____ cl

4) 50,200 cm = _____ dam

5) .0045 kg = _____ mg

Try These: Factor Label Method

6) 2.56 ft = _____ cm

7) 4.3 oz = _____ kg

To do 6 and 7 you need to know the English - Metric conversions:

1 in = _____ cm
1 qt = _____ l
1 lb = _____ g

Now try these. Some of these take several steps:

8) 560 cm = _____ in

9) 1.22 lb = _____ g

10) 4.3 qt = _____ l

11) 2.78 ft = _____ cm

12) 300 cg = _____ lb

13) 3.4 gal = _____ dl

14) 12 oz = _____ mg

15) 10,500 ul = _____ pt

OVER

16) 2 yds = _____ dm

17) 320 lbs = _____ Mg

Now for a challenge:

18) 55 mi/hr = _____ km/hr

19) 15 km/hr = _____ mi/hr

20) 186,000 mi/s = _____ m/hr

Word problems:

21) You have to pour 2 cubic yards (yd^3) of concrete for a patio. What is this volume in cubic meters (m^3) ?

22) If you gained 2 pounds in a week, how many grams is that per day?

23) In gym, you ran the 50 yard dash in 8 seconds. What is your speed in miles per hour?

METRIC SYSTEM

Number	Prefix	Symbol	
1000000	mega	M	
1000	kilo	k	
100	hecto	h	
10	deka	da	
1			meter, liter, gram
.1	deci	d	
.01	centi	c	
.001	milli	m	
.000001	micro	μ	

Length
10mm=1cm
100cm=1 m
1000 m = 1 km

Volume
1000 ml = 1 L
1000 L = 1 KL

$1 mL = 1 cm^3$

$1 L = 1 dm^3 = 1000 cm^3$

Mass
1000 mg = 1 g
1000 g = 1 kg

METRIC TO ENGLISH
1 cm = 0.3937 in = 0.03281 ft
1 m = 39.37 in = 3.281 ft = 1.094 yd
1 km = 3281 ft = 0.6214 mi
1 cm^3 = 0.0610 in^3 = 0.0000353 ft^3
1 L = 1.06 qt = 0.265 gal = 0.0353 ft^3
1 g = 0.0353 oz = 0.00220 lb
1 kg = 2.20 lb = 0.00110 tn
1 metric tn (1000)kg) = 2200 lb = 1.10 tn

1 gal = 4 qts
1 qt = 2 pts
1 pt = 2 cups

ENGLISH TO METRIC
1 in = 2.54 cm = 0.0254 m
1 ft = 30.5 cm = 0.305 m
1 yd = 91.4 cm = 0.914 m
1 mi = 1609 m = 1.609 km
1qt = 946 mL = 0.946 L
1oz = 28350 mg = 28.35 g
1 lb = 453.6 g = 0.4536 kg
1 tn = 907 kg = 0.907 metric tn

12 in = 1 ft
3 ft = 1 yd
1,760 yds = 1 mile
5280 ft = 1 mile

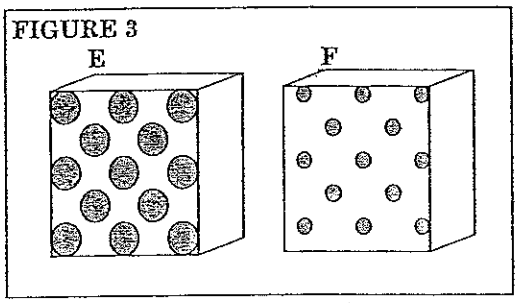
60 sec = 1 min
60 min = 1 hr
24 hrs = 1 day
365 days = 1 year

16 oz = 1 lb

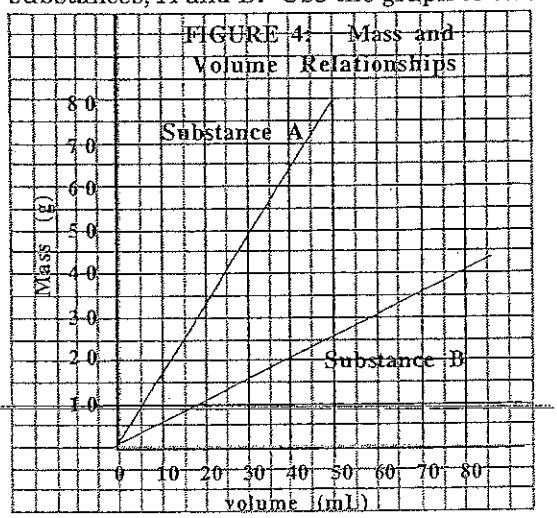
2000 lbs = 1 ton

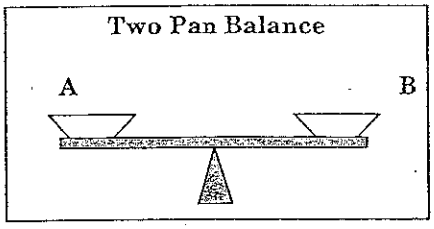
KNOW:
1 in = 2.54 cm
1 qt = .946 L
1 lb = 453.6 g

3. Is object E or object F more dense? [*Assume the particles are uniformly distributed throughout each object, and particles with a larger size have a larger mass.*] Explain your reasoning.



4. In Figure 4 below, a graph shows the relationship between mass and volume for two substances, A and B. Use the graph to answer questions about these two substances.





a) You have built a simple two-pan balance shown abpve to compare the masses of substances A and B. What would happen to the balance if you put **equal masses** of A and B in the two pans? **Equal volumes** of A and B in the two pans? Explain your reasoning.

b) Find the slope of the line for both A and B using correct units. State the physical meaning of the slope for each substance.

c) If you put **10.0 mL of A** in one balance pan, how much **mass of B** would you need in the other pan to make it balance? Explain your reasoning.

d) If you put **35.0 mL of B** in one balance pan, what **volume of A** would you need in the other pan to make it balance? Explain your reasoning.

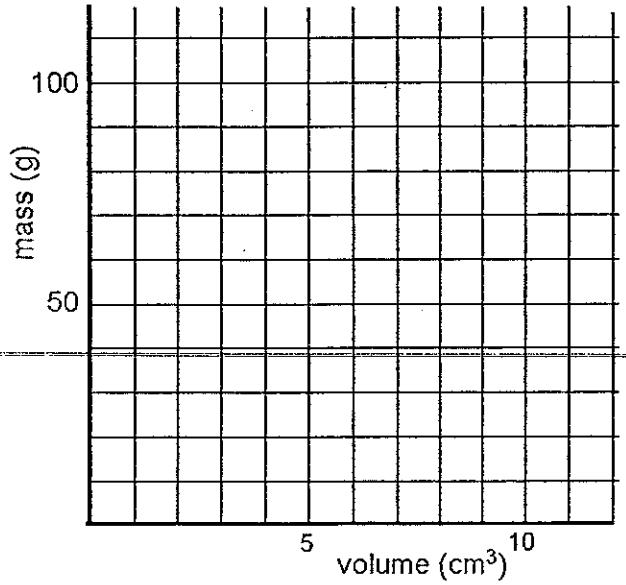
e) Water has a density of 1.00 g/mL. Sketch the line representing water on the graph in Figure 4.

f) Determine whether substance A and B will sink or float when placed in a bucket of water.

A: sink float B: sink float (circle correct response)

Defend your answer using the m-V graph, and your outstanding understanding of density.

Refer to the table of densities at the right to answer the following questions.



Substance	Density (g/mL)
Aluminum	2.70
Titanium	4.54
Zinc	7.13
Tin	7.31
Iron	7.87
Nickel	8.90
Copper	8.96
Silver	10.50
Lead	11.35
Mercury	13.55
Gold	19.30

5. Sketch a graph of mass vs volume for titanium, copper and mercury.

6. You made some cubes out of each metal in the table that each measures 2.00 cm on every side. (all except mercury – why can't you make a cube of mercury?)
   a. What is the volume of each cube in cm^3? in mL? (Show your thinking)

   V = _____ cm^3 V = _____ mL

   b. Find the mass of these metal cubes: (Show your work below)

   lead cube _____

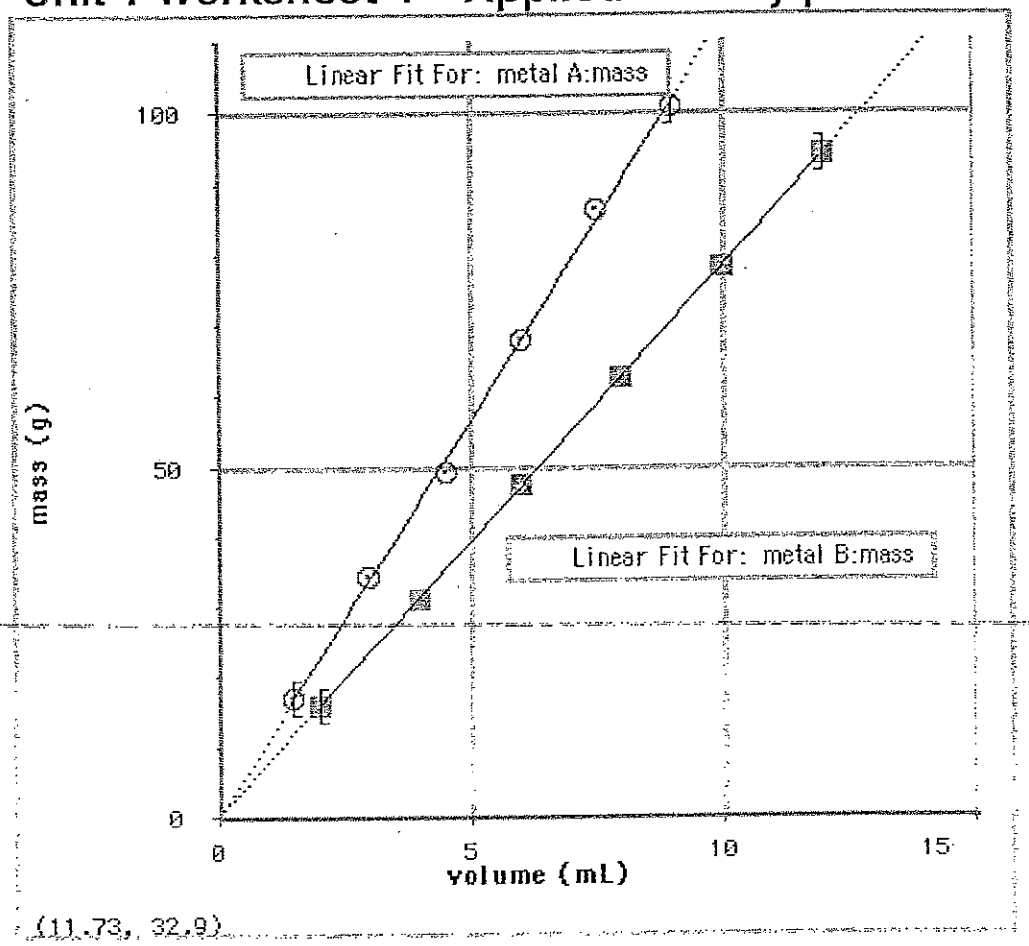
   nickel cube _____

   zinc cube _____

Name_____

Date_____ Pd____

Unit 1 Worksheet 4 – Applied density problems

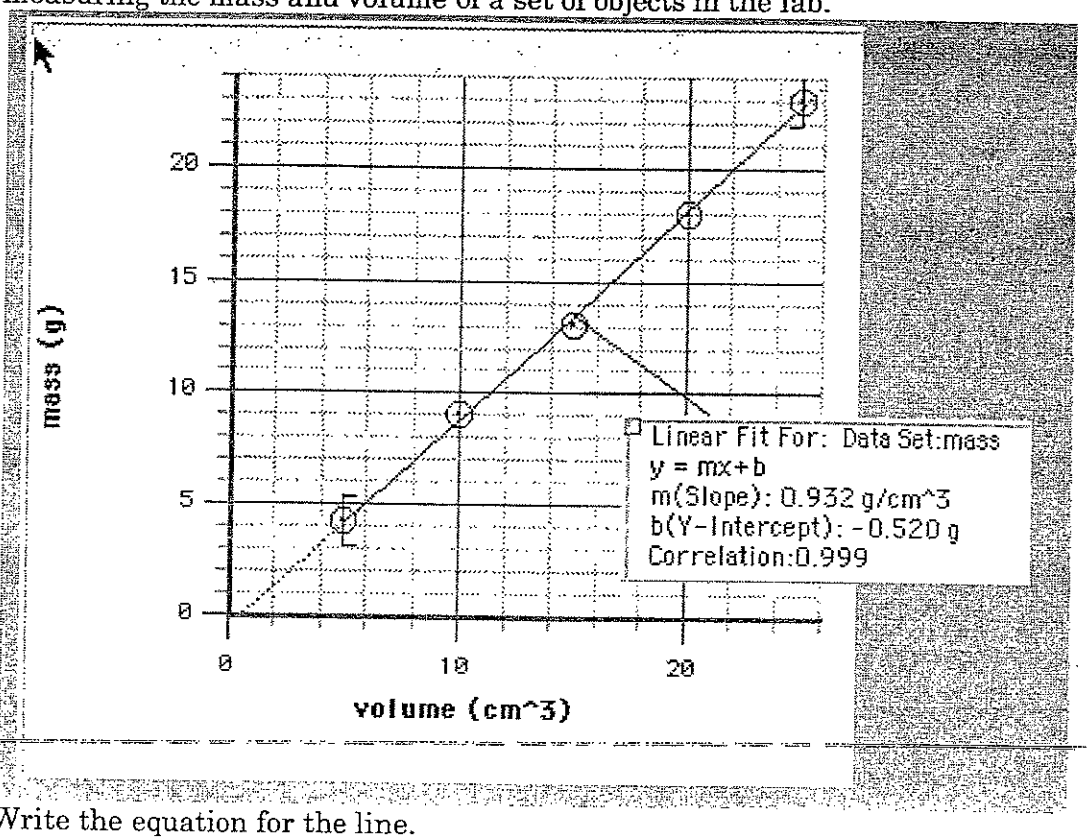


1. Determine the density of each metal. Show all your work and include appropriate units.

2. From the graph, estimate
   a. the mass of 8.0 cm^3 of metal A. _____
   b. the volume of 70 g of metal B. _____
   c. mark on the graph how you found the answers above

3. In the space above right, use the density of B as a factor to determine the answer to 2b. Show the set-up including how the units cancel.

7. The 9th Hr chemistry class produced the following graph when they were measuring the mass and volume of a set of objects in the lab.



a. Write the equation for the line.

b. What information is given by the slope of the graph?

c. Is the y-intercept negligible? Why or why not?

d. What would you predict would happen if you were to put one of the objects in water? Explain.

e. What would you expect to be the mass of a 45 cm^3 piece of the same substance

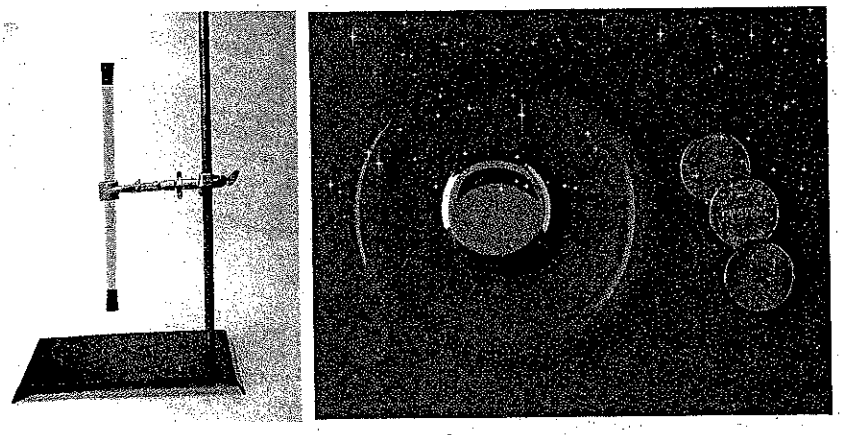


Figure 4-16
Samples of matter in three phases. The brownish gas in the sealed glass tube is nitrogen dioxide. The metal mercury, a liquid, forms a puddle on the surface of the watch glass. Pennies made from an alloy of copper represent a metallic solid.

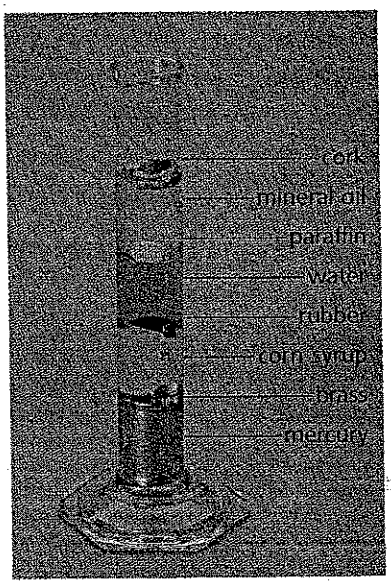
In a sense, ice, liquid water, and steam *are* different substances, because some of their properties are different. However, the differences are entirely physical. Chemically, they are alike. Besides, ice and steam can be changed to liquid water by simply physical means. It is therefore customary to consider water in its three phases to be the same substance. The same applies to any other substance that can exist in more than one phase.

CAN YOU EXPLAIN THIS?

Layers of Liquids

The photo shows a cylinder containing four liquids: mineral oil, water with coloring added to make it easier to see, corn syrup, and mercury. The following solids are floating in the cylinder: pieces of cork, paraffin, hard rubber, and a brass weight. The positions of the layers of liquids and the solid objects will not change visibly from one day to the next if not disturbed.

1. Why do the liquids remain in layers?
2. Why do the solid objects remain in the positions shown?



4-7 Physical and Chemical Properties

Physical properties of a substance are those characteristics that can be observed without the production of new substances. Examples are color, odor, taste, hardness, density, melting and boiling points, and electrical conductivity. Most metals have a set of physical properties that are very different from the properties of nonmetallic substances. Metals are *ductile* (they can be drawn into wires) and *malleable* (they can be hammered into sheets). Also, they have metallic *luster* (they shine in a way typical of metals), and they are good conductors of electricity.

Density Worksheet

1) Determine the volume that 35.2 grams of CCl_4 will occupy if it has a density of 1.60 g/cm^3?

2) The density of ethanol is 0.789 g/cm^3 at 20 C. What is the mass of 150 cm^3 of the solution?

3) A block of lead measures 20.00mm x 3.00 cm x .45 dm. Calculate the mass of this block if the density of lead is 11.34 g/ cm^3.

4) What is the density of a hydrochloric acid solution which has a mass of 17.84 g and occupies 15.00 cm^3?

5) What is the density of a solid piece of iron which has a mass of 11.78 grams and is 30 mm long, 1 cm wide and .05 dm thick?

6) Calculate the density of the following material:

a) 35.0 g of a substance which occupies 25.0 cm^3.

b) 2.75 g of a substance which occupies 250 cm^3.

c) 2.80 kg of a substance which occupies 2.00 L.

7) The mass of 15 cm^3 of a solution is 12 g. Calculate the density of the solution.

8) The density of a solution is 1.5 g/cm^3. Calculate the mass in grams of 10 cm^3 of the solution.

9. The density of a solution is 1.80 g/cm^3. What volume will 360 g of the solution occupy?

10) The first of 3 identical tanks is filled with H_2O, the second with CCl_4, and the third with Hg. The densities of the 3 liquids are 1.00 g/cm^3, 1.58 g/cm^3, and 13.34 g/cm^3 respectively. If the tanks measure 4.00 cm long, 4.00 cm wide and 30.00 mm deep, what is the mass in grams of the contents of each tank?

Density Worksheet

1) The mass of 20.0 cm^3 is 12.0 g. Calculate the density.

2) The density of a solution is 1.35 g/cm^3. Calculate the mass in grams of 10.0 cm^3 of the solution.

3) The density of a solution is 1.83 g/cm^3. What volume will 330. g of the solution occupy?

4) You have 30.0 g of nitric acid (HNO_3). If the density of nitric acid is 1.25 g/ ml, how many ml of nitric acid do you need?

5) You filled a fish tank with water. The fish tank holds 10 gallons. The density of water was 1.00 g/ml. How many pounds of water were used to fill the tank?

6) Which would be heavier? A ball of lead with a diameter of 2.00 cm or a cylinder of iron with a diameter of 3.00 cm and a height of 8.00 cm. (D of lead = 11.3 g/cm^3 and the D of iron = 7.86 g/cm^3)

7) A rectangular rod of copper measured 12.00 mm x 2.00 cm x 1.00 m. If the density of copper is 8.92 g/cm^3, what is the mass in kilograms of the copper rod?

Formulae you may need to know:

$V_{sphere} = 4/3\ \pi\ r^3$

$V_{cylinder} = \pi\ r^2 h$

$V_{rectangle} = l \times w \times h$

$D = m/v$

Honors Chemistry Lab Write-up Format

Your lab write-ups should be done the same way each time. However each lab will be unique. Below you will find the format we will use on all labs. Please read it and use this sheet every time you write up a lab.

1. **Title** - Place the lab number and lab name at the top of a new page in you lab book. If you are using paper, include a cover page.
2. **Purpose** - This is a single statement that states what you will, measure, determine, learn, etc. Example – "To determine the density of lead, steel, brass, aluminum, glass, and rubber."
3. **Method** - This is a very brief description of what is going to take place. It **is not** a procedure. It should be a few simple sentences to let the reader know how the lab was basically done. It should not resemble the procedure you were given to do the lab. Example – "The mass of the objects was measured on the centigram balances. The volume was measured using overflow cans and the water displacement method. The density was then determined mathematically."
4. **Data** – The data format will depend on the lab you are doing. If you are measuring something many times, a table would probably make the most sense. If you are measuring 10 things once each, you will probably want to use a date list. Make sure you put the unit in the heading and on the top number in each column. Do not write the unit on every number in the column. It makes it very difficult to read. Make sure you carry the correct number of significant digits on all numbers. (You teacher will discuss this with you)
5. **Introduction to the Calculations** – In this section you will talk about the theory behind the calculations. It should help the reader understand the calculations. It is not a list of steps that you will follow. Think of it as if you had to tell some other person how the calculations are done. Example – "Density is a comparison of an objects mass and volume. In equation form it would be, $(d = m \div v)$. For each object the measured mass will be divided by the measured volume.
6. **Sample Calculations** - Each calculation should be numbered and contain a label telling the reader what is being calculated. Make sure you include any equation you use, and every number must have a unit. It is called a, sample calculation because you will only show one example of each type of calculation. In the density lab, for instance, you will only show the calculation for one of the metals. The rest are done on scrap paper and the answer is included in the data table. Example - 1. Determine the density of steel by dividing the mass by the volume.
   D = m/v D = 255.14 g/ 32.0 ml D = 7.973125 g/ml d=7.97g/ml
7. **Discussion** - In this section you will convince me that you understand all the concepts in the lab. It is like a written presentation on the topic being studied. Define things, explain things and talk about your answers. Discuss possible errors. You may also have questions on the lab sheet. These questions may refer to the calculations or the discussion. Answer any discussion questions here. Do not answer them in a question section. Incorporate them into you discussion. Your teacher will talk more about this section in class.

8. **Conclusion** - You simply answer the exact question you posed in the purpose. In the case of the density lab you would put, "This lab showed that the density of steel is _____, brass_____ etc. (State your value)

Just a few more comments:

- You must use pen in labs
- If you need to change anything, you put a single line through it if it is a word or sentence, or a large "X" if a large section is not wanted. It must still be able to be read under your marks.
- Never use white out under any circumstances
- Never use the words, "I, we, you" in lab. Write it as though you are telling someone how to do things. Example - "heat a beaker with the Bunsen burner" or "Error in the lab could have occurred when the contents of the beaker were poured into another container because it was difficult to remove all the liquid"
- These labs should be written-up totally on your own. NO part of your lab should be like your partners lab except the data section. If you have a lab that is too similar to another person's lab, both labs will receive a grade of zero.
- These labs are a significant part of you grade. You cannot write a good lab report at the last minute in a short period of time. You will be given two to three days to write the lab up.

Chemistry Lab #1 Density

In this lab you will determine the density of a number of different substances by measuring their mass and volume. The volume will be measured using water displacement.

Procedure:

1) Obtain a triple beam balance and a centigram balance. You are going to use both balances in order to learn how to use them correctly.
2) You are going to measure the density of the following substances: lead, steel, brass, aluminum, glass, and rubber. Obtain one or two of these substances. Do not get all six at one time since there are not enough for each group to get a complete set. Once you have finished with the ones you have, return them and get two more.
3) The triple beam balance will be used to measure the lead, brass and steel. The centigram balance will be used to take the mass of the aluminum, glass and rubber. Place one of your substances on the appropriate balance and find its mass. Record this mass in your data table.
4) Now find the volume of the same substance you found the mass for in number 3. You will do this using an overflow can. Fill the can to the top so that water flows out the spout into the sink. When the water stops, hold a 50 ml graduated cylinder under the spout. Now, as gently as possible, lower the object into the water. The water you catch will equal the volume of the object. Record the volume in your data table.
5) Repeat steps 3 and 4 until you have recorded the mass and volume of all six substances. Return both balances to the cabinet. Your teacher will show you how they are arranged.

Discussion questions:

1) Calculate the density of all six substances and record the value in your data table.
2) Compare your values with the actual values by calculating a percent error. Record this percent in your data table. (Your teacher will give you the known values in class)
3) What does the density of a substance tell you about that substance?
4) What substances tend to have high densities? What substances have low densities?
5) Using density, describe when one substance will float on another.

SUBSTANCE	MASS	VOLUME	DENSITY	THEORETICAL DENSITY	% ERR
Lead					
Steel					
Brass					
Aluminum					
Glass					
Rubber					

DENSITY LAB

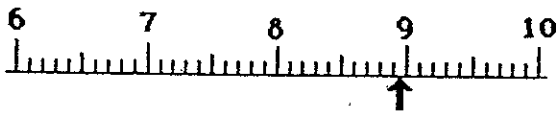
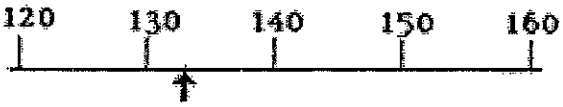
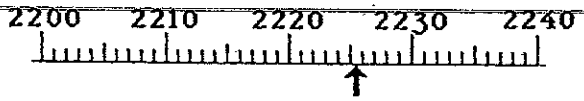
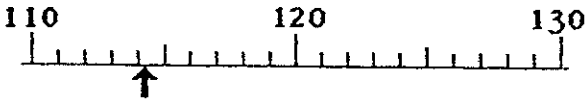
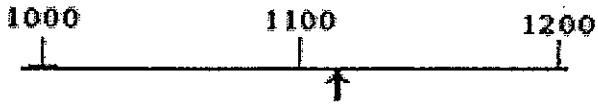
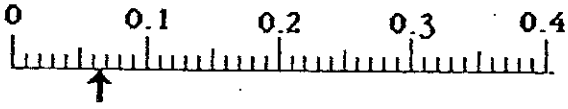
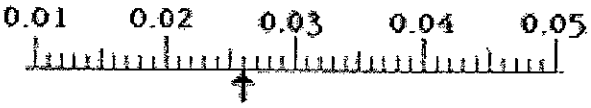
GROUP	m(g)	V(ml)	D(exp)	D(theor)	%error
	Pb				
	Steel				
	Brass				
	Al				
	Glass				
	Rubber				
	Pb				
	Steel				
	Brass				
	Al				
	Glass				
	Rubber				
	Pb				
	Steel				
	Brass				
	Al				
	Glass				
	Rubber				
	Pb				
	Steel				
	Brass				
	Al				
	Glass				
	Rubber				

Name _____

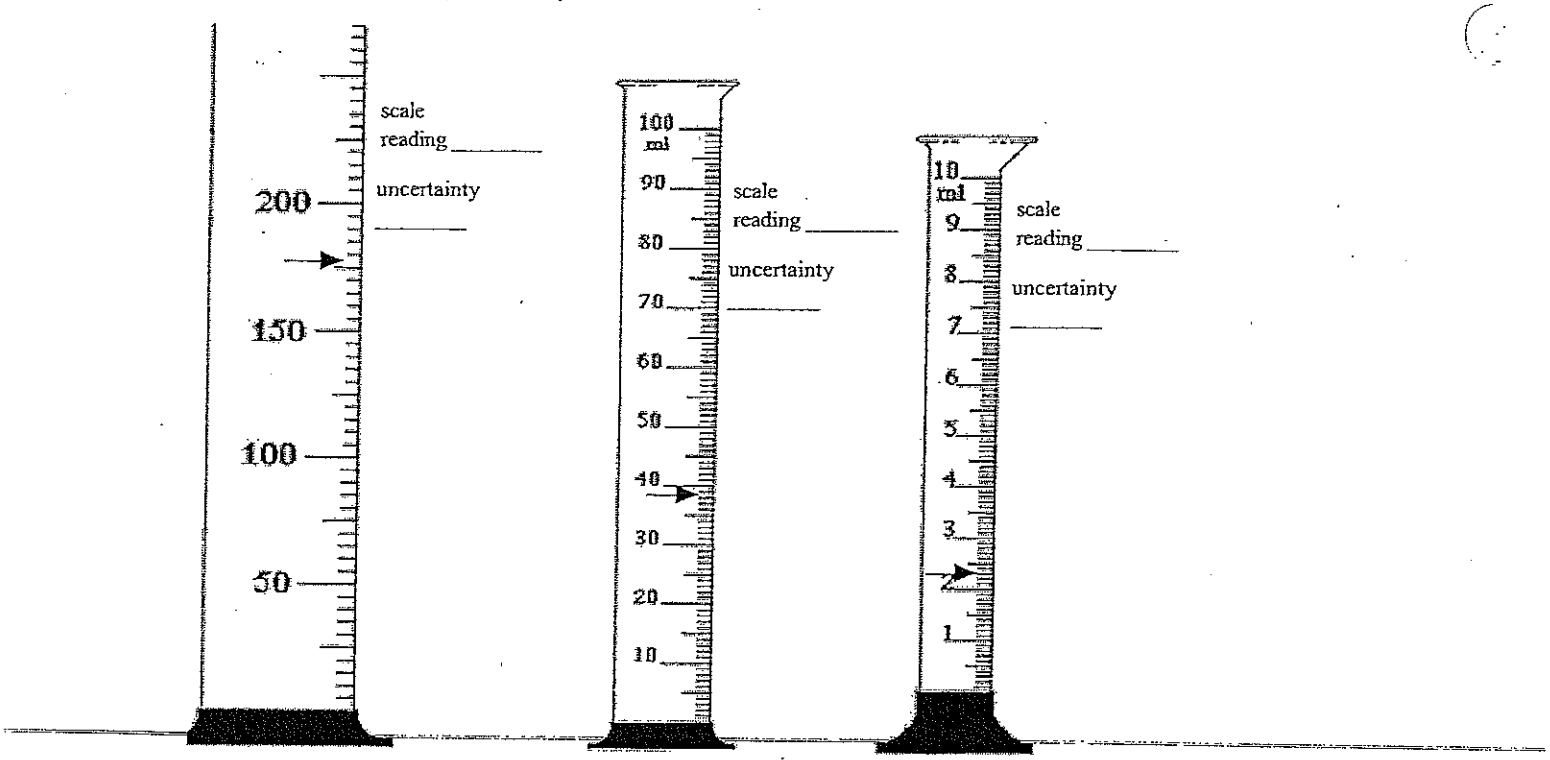
Date _____ Pd ____

Unit 1: Worksheet 2 - Reading Scales

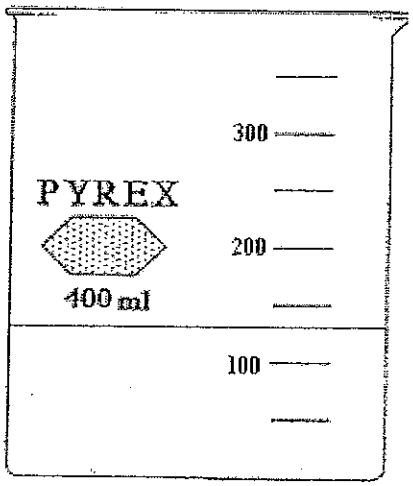
For each of the following, write the scale reading, then the number of significant figures in the reading.

		Reading	SF's
1.		_____	_____
2.		_____	_____
3.		_____	_____
4.		_____	_____
5.		_____	_____
6.		_____	_____
7.		_____	_____
8.		_____	_____

For each of the volume devices below record the scale reading and indicate the uncertainty in the measurement (± _____).

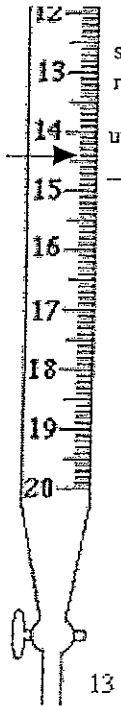


9 10 11



12 scale reading _____

uncertainty _____



scale reading _____

uncertainty _____

13

CHEMISTRY WORKSHEET: MEASUREMENTS, UNCERTAINTY AND SIGNIFICANT DIGIT

No measurement is absolutely accurate due to the problem of "uncertainty." There are (1) measuring limitations of the equipment, and (2) differences in the skills of the observers. With practice, measuring skills can be developed, and an understanding of the limitations of the equipment can be learned.

The laboratory equipment:

1. What is the smallest division (value in grams) on our balance? _____

   (a) Record the weight of a 250 ml beaker _____

   (b) Record the weight of a #4 rubber stopper _____

   (c) Record the weight of a pipe stem (clay) triangle _____

   (d) Record the weight of an evaporating dish _____

   (e) Record the weight of a watch glass _____

   (f) Record the weight of a stirring rod _____

2. What is the smallest division of a 250 ml flask? _____

   (a) Add a small amount of water, record volume _____

   (b) Repeat above _____

   (c) Repeat again _____
   *Note: estimate the last digit

3. What is the smallest division on a 10 ml graduated cylinder? _____

   (a) Add a small amount of water, record volume _____

   (b) Repeat above _____

   (c) Repeat again _____
   *Note: estimate the last digit

4. What is the smallest division on a 100 ml graduated cylinder? _____

   (a) Add a small amount of water, record volume _____

   (b) Repeat above _____

   (c) Repeat again _____
   *Note: estimate the last digit

5. What is the smallest division on a thermometer? _____

(a) Record room temperature _____

(b) Record body temperature by carefully holding your hand around the bulb of the thermometer. What is your body temperature in degrees F? _____

RECORDING EXPERIMENTAL MEASUREMENTS USING THE SIGNIFICANT DIGIT METHOD

All measurements are reported by recording a number which includes the digits which are shown as actual divisions plus one estimated digit.

(a) Assume the following are experimentally measured quantities. List the number of significant digits appearing in each number below:

1.	15 cm _____	7.	0.0602 g _____
2.	0.15 cm _____	8.	6.050 g _____
3.	15.0 cm _____	9.	3.00 g _____
4.	0.015 cm _____	10.	0.003567 g _____
5.	1.250 cm _____	11.	9.0900 g _____
6.	13.002 cm _____	12.	1.01 ℓ _____

Significant Figures

Sig figs are an essential part of all laboratory measurements and calculations. They indicate the precision of the measurement. If you use too many significant figures, this gives the impression that your measuring instrument was of higher quality than it really was.

Many people think that using a lot of numbers indicates that you have been very careful and have done a very accurate job. To experienced laboratory scientists using too many numbers just indicates that you really don't understand the instruments that you are using. For example, in labs you can use many instruments for measuring volume. A graduated cylinder, a pipette, an erlenmeyer flask etc. The number of calibrations on the measuring device limits the number of digits that you can use in your answer. Because the pipette can measure 0.01 ml , you could report your answer to 0.01 ml. However, the erlenmeyer flask has 25 ml increments. How can you report measurements using this device ?

Using significant figures, a measurement should consist of all the digits that you can read and one estimated digit. Significant figure measurements consist of certain digits and a final digit that it an estimate.

Remember: All measurements consist of digits and UNITS !

Which numbers are significant ?	Examples
1. All nonzero numbers are significant.	93 = 2 sig fig
2. Zero's between nonzero numbers are significant	103 = 3 sig fig
3. Place holder zeros are not significant.	0.03 = 1 sig fig
Use scientific notation to reduce ambiguity	1200= 2 sig fig
4. Scientific notation , all numbers and zeros are	40,000 = 1 sig fig
significant. (not the x 10 however)	4.0×10^4=2 sig fig
	4.000×10^4=4 sig fig
5. Zeros after a decimal are significant if the number is greater than one.	30.0=3 sig fig
6. Adding a decimal makes a zero in the units column significant.	30. = 2 sig fig

Rounding using Sig Figs:

If the digit immediately to the right of the last significant digit you want to keep is

1. Greater than 5 .....then you should increase the last significant digit

2. Less than 5 ..... then the last significant digit should stay the same

3. 5 followed by ..... then increase the last sig fig by 1 nonzero digit

4. 5 not followed by nonzero ..... then increase last sig fig by 1 digit and preceded by an odd digit

5. 5 not followed by nonzero..... then the last sig fig stays the same digit, and preceded by an even digit

Practice Problems:

How many significant figures are contained in each of the following measurements? Which sig fig rule applies to each?

a. 37.4 mm _____Rule_____

b. 6070 dm_____Rule_____

c. 0.00903 km_____ Rule_____

d. 350.0 K_____Rule_____

e. 450 m_____ Rule_____

f. 0.8030L _____Rule _____

Round off each of the following measurements to the indicated number of significant figures.

a. 35.27 g to 3 sig fig
b. 0.414 kL to 2 sig fig
c. 87.257 dm to 3 sig fig
d. 1.35K to 2 sig fig
e. 6250 cm to 2 sig fig
f. 4.681 cm to 2 sig fig

SIGNIFICANT FIGURES

Significant digits are used in chemistry when referring to measurements or data. Because instruments used to obtain measurements have limitations due to their quality or design . All measurements include certain and uncertain digits. The uncertain digit is the last number . You are sure of the certain digits. The last digit is the uncertain digits. You are unsure of this one because of instrument limitations.

All data and calculations that you do in this class will involve significant digits.

What are significant digits ?

All integers (numbers 1 through 9) are always significant.

Sometimes zero is also significant.

Zeros between integers are always significant ex: 507 has three significant figures.

Zeros before integers are not significant . Ex : 00.0507 has three significant figures.

Zeros after the last integer are significant ONLY when there is a decimal point in the number . Ex: 00.05070 has four significant digits. 5070 has three significant digits.

When using scientific notation , all zeros and integers are significant.

ex: 5.7×10^3 has two significant digits 5.70×10^3 has three significant digits.

ADDITION AND SUBTRACTION

The measurement with the fewest decimal places determines the correct number of significant digits in your answer.

Ex: 7.83 + 5.4 = 13.2 (NOT 13.23)

MULTIPLYING AND DIVIDING

The measurement with the fewest significant digits determines the number of significant digits in the answer.

Ex: 0.59 - 1.01 = 0.58 NOT 0.584158415 because 0.59 has only two significant digits your answer can have only two significant digits.

The answer must never have more significant digits than the original numbers !

Worksheet: Scientific Notation and Significant Figures

Write in scientific notation:

1) 100,000
2) 10,000,000
3) 0.01
4) 0.0000001
5) 400,000,000,000
6) 50,000
7) 0.0088
8) 0.0000045
9) 38,200,000
10) 280,400,000

Write as a whole number or decimal:

1) 1.00×10^{4}
2) 1.0×10^{-4}
3) 6×10^{6}
4) 5.2×10^{-4}
5) 3.82×10^{2}
6) 1.687×10^{-5}
7) 16.85×10^{-3}
8) 84.3×10^{6}
9) 9.6503×10^{-4}
10) 6.854×10^{14}

Do the following calculations and put the answer in scientific notation and correct number of significant digits.

1) $0.08990 + 52 =$
2) $0.9 - 0.00005 =$
3) $63 + 93 =$
4) $4 + 6 =$
5) $0.06 + 2 =$
6) $(9.43 \times 10^{5}) + (8.82 \times 10^{4}) =$
7) $(2.78 \times 10^{-5}) - (6.12 \times 10^{-6}) =$
8) $(1.6 \times 10^{2})(2.4 \times 10^{-3}) =$
9) $(75.4 \times 10^{2})(0.774 \times 10^{-2}) / (3.4 \times 10^{3}) =$

Scientific Notation worksheet

Convert the following into scientific notation

1. 785

2. 10,975

3. 387,631

4. 0.0155

5. 0.00068

6. 0.00000798

Convert the following into whole numbers or decimals

7. 1.85×10^{2}

8. 2.63×10^{-2}

9. 7.49×10^{4}

10. 6.33×10^{6}

11. 3.57×10^{-5}

12. 2.99×10^{-8}

Solve the following problems and express your answer in scientific notation.

13. $(1.7 \times 10^{2}) + (2.1 \times 10^{3})$

14. $(3.85 \times 10^{3}) + (6.52 \times 10^{-2})$

15. $(6.94 \times 10^{2}) - (3.75 \times 10^{4})$

16. $(7.573 \times 10^{-3}) - (8.48 \times 10^{-5})$

17. $(4.6 \times 10^{4}) \times (1.8 \times 10^{2})$

18. $(7.874 \times 10^{-6}) \times (6.38 \times 10^{4})$

19. $(5.25 \times 10^{3}) \div (1.6 \times 10^{2})$

20. $(6.32 \times 10^{6}) \div (1.153 \times 10^{-2})$

85

Worksheet: Scientific Notation and Significant Figures

Write in scientific notation:
1) 100,000
2) 10,000,000
3) 0.01
4) 0.0000001
5) 400,000,000,000
6) 50,000
7) 0.0088
8) 0.0000045
9) 38,200,000
10) 280,400,000

Write as a whole number or decimal:

1) 1.00×10^{4}
2) 1.0×10^{-4}
3) 6×10^{6}
4) 5.2×10^{-4}
5) 3.82×10^{2}
6) 1.687×10^{-5}
7) 16.85×10^{-3}
8) 84.3×10^{6}
9) 9.6503×10^{-4}
10) 6.854×10^{14}

Do the following calculations and put the answer in scientific notation and correct number of significant digits.

1) $0.08990 + 52 =$
2) $0.9 - 0.00005 =$
3) $63 + 93 =$
4) $4 + 6 =$
5) $0.06 + 2 =$
6) $(9.43 \times 10^{5}) + (8.82 \times 10^{4}) =$
7) $(2.78 \times 10^{-5}) - (6.12 \times 10^{-6}) =$
8) $(1.6 \times 10^{2})(2.4 \times 10^{-3}) =$
9) $(75.4 \times 10^{2})(0.774 \times 10^{-2}) / (3.4 \times 10^{3}) =$

Accel

Put answer in scientific notation & take into consideration significant digits

III. Carry out the following operations, using exponential notation:

1. 6000 x 0.004 =
2. 50 x 0.0000002 =
3. 800 ÷ 40 =
4. 4000 + 80 + 500 =
5. 30,000 x 450,000 =
6. 0.006 ÷ 0.2 =
7. 0.0002 x 0.04 =
8. 32,000 ÷ 400,000 =
9. 1,000,000 x 7900 =
10. 60,000 + 7000 + 80,000 =
11. 0.000045 ÷ 0.009 =
12. 0.0024 x 20,000 =
13. 0.0064 ÷ 8000 =
14. 0.090 ÷ 0.00003 =
15. 0.0004 x 5000 x 0.02 =
16. 450 + 250 + 7500 =
17. 400 ÷ 200,000 =
18. 6 ÷ 30,000 =
19. 90 x 0.02 x 0.00001 =
20. 0.000081 ÷ 90,000 =
21. 40,000,000 x 0.001 x 0.20 =
22. 300,000,000 ÷ 0.0006 =
23. 6450 + 28,200 + 3740 =
24. 0.060 x 20,000 x 0.0002 =
25. 4 ÷ 0.0000002 =
26. 13 x 0.000001 x 4000 =
27. 4850 - 370 =
28. 0.006 x 0.0002 x 0.02 =
29. 0.0084 ÷ 42 =
30. 1400 x 2,000,000 x 0.004 =
31. 0.0045 + 0.0550 + 0.0015 =
32. 0.00072 ÷ 900,000 =
33. 0.0005 x 200 x 40,000,000 =
34. 0.049 – 0.00017 =

Honors Chemistry Homework Significant Digits

1. How many significant digits do each of the following numbers have?
   a. 1.00060 _____ b. .000234 _____ c. .10000 _____
   d. 100000 _____ e. 100100 _____ f. 1.3400×10^4 _____

2. Round off the numbers to 3 significant digits.
   a. 1.234234 _____ b. 48.965_____ c. .000023000 _____

3. Do the following calculations and round off the answer to the correct number of significant digits.

a. (5.78)(1.00876)(.0024) = _____ d.. (20)9876)(.000025) ÷ (8790)53.60) = _____

b. (9.760)10500) ÷ (4.300) = _____ e. (2.909) – (1.34) = _____

c. 209.3 + 34 + .0005 + 28 = _____ f. 3.9820 -1.88 + 2930 = _____

4. Do the following problems and round the answer off to the correct number of significant digits.

   a. An object has a mass of 234.86g and a volume of 26.2 ml. What is the density?

   b. How much would a 2.00 liter bottle weigh if it were filled with mercury? (d = 13.6 g/ml for Hg)

   c. What volume would 500 grams of aluminum have if the density is 2.712 g/ml?

Chemistry Review Sheet

Students entering chemistry often have different backgrounds and levels of skills and knowledge. As we start our study, we need to have some common ground. On this sheet you will find things that we are going to assume you know and understand. If that is not the case, we will need to take steps to correct that. (Please don't panic if you are unfamiliar with some items. We will fix that.)

1. Scientific Notation - You should be able to do all mathematical operations with numbers in scientific notation. You should have a good feeling for how big or small a number is based on the exponent.
2. Conversions using the factor label method – You should be able to convert from one unit to another using the factor label method given the correct conversion factor. (Note – it is the setup that is important. I know metric conversions can be done in your head by moving the decimal point. I would encourage you to do so to save time. But, we use this for so many other conversions in chemistry that the **process** is what you must master.) On the problems below, I need to see the conversion fractions, not necessarily the answer.
3. Solving simple algebraic equations – You should be able to manipulate equations to isolate the variable. This should also include the manipulation of the units.
4. Metric relationships – You must know what the prefixes represent and how they are related. (Only the common ones we use)
5. Graphing - You should be able to draw a line graph from the data by hand and by computer. More important, however, is that you can interpret what the graph represents.

1. $(3.857 \times 10^{4})(5.43 \times 10^{-9}) = ?$ 2. $(1.23 \times 10^{5})(3.998 \times 10^{-8}) / (7.77 \times 10^{-15})(97)$

3. 4.874 m = __?__ cm 4. 15000 ml = __?__ L 5. 4567000 cm^3 = __?__ m^3

6. 55miles/hr = __?__ m/sec (5280ft = 1 mile, 2.54cm = 1inch) 7. 4 days = __?__ seconds

8. – 9. Use the equation, $q = mC\Delta T$ 8. q = 5400 cal , m = 139 g , C = .031 cal/g°C find Δt

9. m = .038 Kg, C = .107 cal / g°C, Δt = 30.7 °C find q.

10 – 12. Use the equation for density, d = m/v

10. A substance has a density of 4.34 g/cm^3 and a mass of 4.39g . What is the volume?
11. What is the density of a substance that has a mass of 500 g in a volume of 28 ml?
12. What mass is contained in 45 ml of a substance with a density of 2.25g/ L?